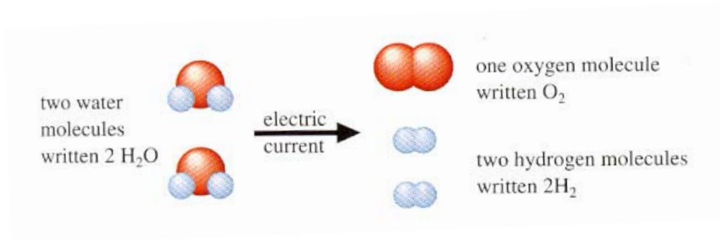
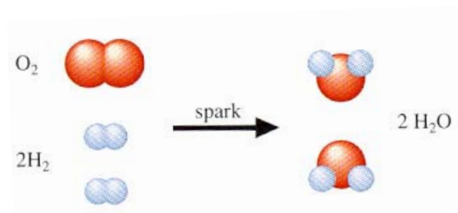
AP* Chemistry
CHEMICAL FOUNDATIONS

1.1 Chemistry: An Overview

- **<u>Matter</u>** – takes up space, has mass, exhibits inertia
  - composed of atoms only 100 or so different types
  - water made up of one oxygen and two hydrogen atoms
  - Pass an electric current through it to separate the two types of atoms and they rearrange to become two different types of molecules



- reactions are reversible

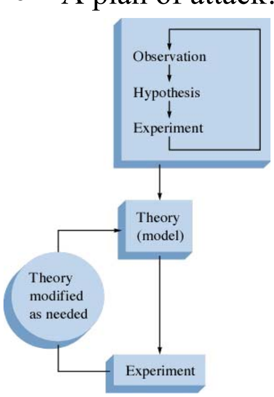


<u>Chemistry</u> – the study of matter and energy and more importantly, the changes between them

- **<u>Why study chemistry?</u>**
  - become a better problem solver in <u>all</u> areas of your life
  - safety – had the Roman's understood lead poisoning, their civilization would not have fallen
  - to better understand all areas of science

1.2 The Scientific Method

- A plan of attack!



The fundamental steps of the scientific method.

Steps in the Scientific Method

1. *Making observations.* Observations may be *qualitative* (the sky is blue; water is a liquid) or *quantitative* (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a **measurement**) involves both a number and a unit.
2. *Formulating hypotheses.* A **hypothesis** is a *possible* explanation for an observation.
3. *Performing experiments.* An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether or not the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

- Repetition of experiments is key

- **Theory** – hypotheses are assembled in an attempt at *explaining* "why" the "what" happened.
- **Model** – we use many models to explain natural phenomenon – when new evidence is found, the model changes!

Robert Boyle

- **Robert Boyle**
  - love to experiment with air
  - created the first vacuum pump
    - coin and feather fell at the same rate due to gravity in a vacuum since there is no air resistance.
  - $P_1V_1 = P_2V_2$
  - defined elements as anything that cannot be broken down into simpler substances

- **Scientific Laws** – a summary of observed (measurable) behavior [a theory is an explanation of behavior]

  A law summarizes what happens; a theory (model) is an attempt to explain WHY it happens.
  - **Law of Conservation of Mass** – mass reactants = mass products
  - **Law of Conservation of Energy** – (a.k.a. first law of thermodynamics)

    Energy CANNOT be created NOR destroyed; can only change forms.
  - scientists are human and subjected to
    - Data misinterpretations
    - Emotional attachments to theories
    - Loss of objectivity
    - Politics
    - Ego
    - Profit motives
    - Fads
    - Wars
    - Religious beliefs
- **Galileo** – forced to recant his astronomical observations in the face of strong religious resistance.
- **Lavoisier** – "father of modern chemistry"; beheaded due to political affiliations.
- The need for better explosives; (rapid change of solid or liquid to gas where molecules become ≈2,000 diameters farther apart and exert massive forces as a result) for wars have led to
  -fertilizers that utilizes nitrogen
  - nuclear devices

1.3 Units of Measurement

A quantitative observation, or measurement, ALWAYS consists of two parts: a *number* and a *unit*. Two major measurements systems exist: English (US and some of Africa) and Metric (the rest of the globe!)

- SI system – 1960 an international agreement was reached to set up a system of units so scientists everywhere could better communicate measurements. Le Système International in French; all based upon or derived from the metric system

Table 1.1 The Fundamental SI Units

Physical Quantity	*Name of Unit*	*Abbreviation*
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

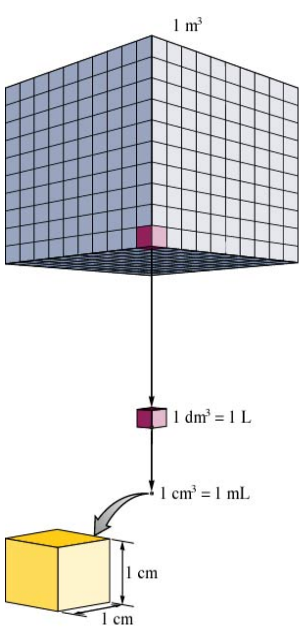
Table 1.3 Some Examples of Commonly Used Units

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

Table 1.2 The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	*Symbol*	*Meaning*	*Exponential Notation**
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.



KNOW THESE UNITS AND PREFIXES!!!

- <u>Volume</u> – derived from length
  consider a cube 1m on each edge $\therefore$ 1.0 m^3
  - a decimeter is 1/10 of a meter so
  $(1m)^3 = (10dm)^3 = 1{,}000\ dm^3$
  $1dm^3$ = 1 liter (L) and is slightly larger than a quart also
  $1dm^3 = 1\ L = (10\ cm)^3 = 1{,}000\ cm = 1{,}000\ mL$
  SINCE
  $1\ cm^3 = 1\ mL = 1$ gram of H_2O (at 4ºC if you want to be picky)

- <u>Mass vs. Weight</u> – chemists are quite guilty of using these terms interchangeably.
  - <u>mass</u> (g or kg) – a measure of the resistance of an object to a change in its state of motion (ie. exhibits inertia); the quantity of matter present.

  - <u>weight</u> (a force $\therefore$Newtons) – the response of mass to gravity; since all of our measurements will be made here on Earth, considered the acceleration due to gravity a constant so we'll use the terms interchangeably as well *although* it is incorrect.

We "weigh" chemical quantities on a **balance** **<u>NOT</u> a scale!!**

- **gravity** – varies with altitude here on planet Earth
    - the closer you are to the center of the Earth, the stronger the gravitational field SINCE it originates from the center of the Earth.
    - Every object has a gravitational field – as long as you're on Earth, they are masked since the Earth's field is so HUGE compared to the object's.
    - The strength of the gravitational field $\propto$ mass
    - Ever seen astronauts in space that are "weightless" since they are very far removed from the center of Earth? Notice how they are constantly "drawn" to the sides of the ship and must push away?
    - The ships' mass is greater than the astronaut's mass $\therefore$ "g" is greater for the ship and the astronaut is attracted to the ship just as you are attracted to Earth!
    - The moon has $\frac{1}{6}$ the mass of the Earth $\therefore$ you would experience $\frac{1}{6}$ the gravitational field you experience on Earth and $\therefore$ you'd WEIGH $\frac{1}{6}$ of what you weigh on Earth.

Physics moment:

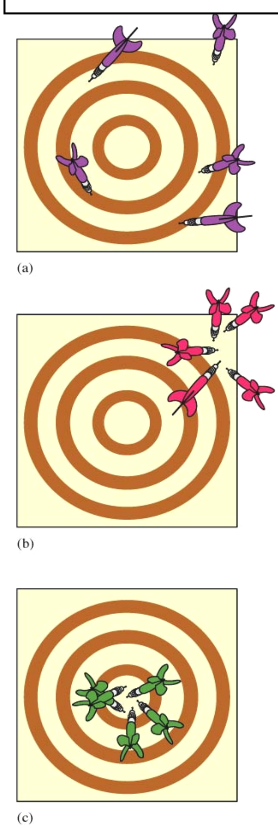
$F_w = ma$
$F_w = mg$

$$F_w = m\left(\frac{9.8m}{s^2}\right)$$

$\therefore$ its units are

$$N = (kg)\left(\frac{m}{s^2}\right)$$

- **Precision and Accuracy**
    - **accuracy** – correctness; agreement of a measurement with the true value
    - **precision** – reproducibility; degree of agreement among several measurements.
    - **random or indeterminate error** – equal probability of a measurement being high or low
    - **systematic or determinate error** – occurs in the same direction each time



The results of several dart throws show the difference between precise and accurate. (a) Neither accurate nor precise (large random errors). (b) Precise but not accurate (small random errors, large systematic error). (c) Bull's-eye! Both precise and accurate (small random errors, no systematic error).

Exercise 1.2 Precision and Accuracy

To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret and then read the volume delivered. Following are the results of five trials:

Trial	*Volume Shown by Graduated Cylinder*	*Volume Shown by the Buret*
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	*25 mL*	*26.54 mL*

Is the graduated cylinder accurate?

Note that the average value measured using the buret is significantly different from 25 mL. Thus this graduated cylinder is not very accurate. It produces a systematic error (in this case, the indicated result is low for each measurement).

1.5 Significant Figures and Calculations

Rules

- Non zero digits are significant
- A zero is significant if it is
  - "terminating AND right" of the decimal [must be both]
  - "sandwiched" between significant figures
- Exact or counting numbers have an ∞ amount of significant figures as do fundamental constants

Exercise 1.3 Significant Figures

Give the number of significant figures for each of the following results.

a. A student's extraction procedure on tea yields 0.0105 g of caffeine.

b. A chemist records a mass of 0.050080 g in an analysis.

c. In an experiment, a span of time is determined to be 8.050×10^{-3} s .

a. three
b. five
c. four

Rules for calculating

- × and ÷ The term with the least number of *significant figures* (∴ least accurate measurement) determines the number of significant figures in the answer.

$$4.56 \times \underline{1}.\underline{4} = 6.38 \xrightarrow{\text{corrected}} \underline{6}.\underline{4}$$

- + and (−) The term with the least number of *decimal places* (∴ least accurate measurement) determines the number of significant figures in the answer.

$$\begin{array}{ll} 12.11 & \\ 18.0 & \leftarrow \text{limiting term} \\ \underline{1.013} & \\ 31.123 & \xrightarrow{\text{corrected}} 31.1 \end{array}$$

- pH – the *number of significant figures in least accurate measurement* determines *number decimal places* on the reported pH

Rounding Rules:

- Round at the end of all calculations
- Look at the significant figure one place beyond your desired number of significant figures if > 5 round up; < 5 drop the digit.
- Don't "double round" 4.348 to 2 SF = 4.3 NOT the 8 makes the 4 a 5 then 4.4. [Even though you may have conned an English teacher into this before!]

1.6 Dimensional Analysis

Table 1.4 English–Metric Equivalents

Length	1 m = 1.094 yd
	2.54 cm = 1 in
Mass	1 kg = 2.205 lb
	453.6 g = 1 lb
Volume	1 L = 1.06 qt
	1 ft^3 = 28.32 L

Consider a pin measuring 2.85 cm in length.
What is its length in inches?

2.54 cm = 1 inch $\therefore$ you can write 2 Conversion factors: $\dfrac{1 \text{ in}}{2.54 \text{ cm}}$ or $\dfrac{2.54 \text{ cm}}{1 \text{ in}}$

To convert multiply your quantity by a conversion factor that "cancels" the undesirable unit and puts the desired unit in the numerator.

$$2.85\ \cancel{\text{cm}} \times \frac{1 \text{ in}}{2.54\ \cancel{\text{cm}}} = 1.12 \text{ in}$$

Exercise 1.5 Unit Conversions I
A pencil is 7.00 in. long. What is its length in centimeters?

17.8 cm

Exercise 1.6 Unit Conversions II
You want to order a bicycle with a 25.5-in. frame, but the sizes in the catalog are given only in centimeters. What size should you order?

64.8 in

Exercise 1.7 Unit Conversions III
A student has entered a 10.0-km run. How long is the run in miles?

We have kilometers, which we want to change to miles. We can do this by the following route:

kilometers $\rightarrow$ meters $\rightarrow$ yards $\rightarrow$ miles

To proceed in this way, we need the following equivalence statements:

1 km = 1000 m
1 m = 1.094 yd
1760 yd = 1 mi

= 6.22 mi

Exercise 1.8 Unit Conversions IV
The speed limit on many highways in the United States is 55 mi/h. What number would be posted in kilometers per hour?

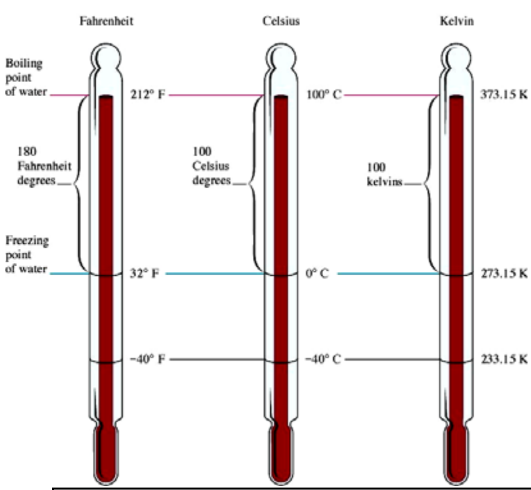
88 km / h

Exercise 1.9 Unit Conversions V
A Japanese car is advertised as having a gas mileage of 15 km/L. Convert this rating to miles per gallon.

35 mi / gal

1.7 Temperature

Three Scales



Notice a degree C = a degree K

$$T_F = T_C \times \frac{9°F}{5°C} + 32°F$$

$$T_K = T_C + 273.15\ K$$

$$T_C = T_K - 273.15\ °C$$

Exercise 1.10 Temperature Conversions I
Normal body temperature is 98.6°F. Convert this temperature to the Celsius and Kelvin scales.

98.6°F = 37.0°C
98.6°F = 310.2 K

Exercise 1.11 Temperature Conversions II

One interesting feature of the Celsius and Fahrenheit scales is that –40°C and –40°F represent the same temperature. Verify that this is true.

Exercise 1.12 Temperature Conversions III

Liquid nitrogen, which is often used as a coolant for low-temperature experiments, has a boiling point of 77 K. What is this temperature on the Fahrenheit scale?

T_F = –319 ° F

1.8 Density

$$\textbf{Density} = \frac{\text{mass}}{\text{volume}}$$

Exercise 1.13 Determining Density

A chemist, trying to identify the main component of a compact disc cleaning fluid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20°C. The following are the names and densities of the compounds that might be the main component.

Compound	Density $\left(\frac{g}{cm^3}\right)$ at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

Which of these compounds is the most likely to be the main component of the compact disc cleaner?

Density = 0.7850 g / cm^3 $\therefore$ isopropyl alcohol

1.9 Classification of Matter

States of Matter

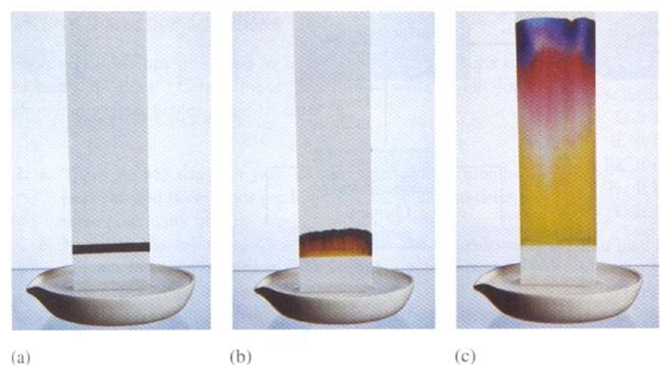
- **<u>solid</u>** – rigid; definite shape and volume; *molecules close together vibrating about fixed points ∴ virtually incompressible*
- **<u>liquid</u>** – definite volume but takes on the shape of the container; *molecules still vibrate but also have rotational and translational motion and can slide past one another BUT are still close together ∴ slightly compressible*
- **<u>gas</u>** – no definite volume and takes on the shape of the container; *molecules vibrate, rotate and translate and are independent of each other ∴ VERY far apart ∴ highly compressible*
  - **<u>vapor</u>** – the gas phase of a substance that is normally a solid or liquid at room temperature
  - **<u>fluid</u>** – that which can flow; gases and liquids

- **<u>Mixtures</u>** – can be **physically** separated
  - **<u>homogeneous</u>** – have visibly <u>in</u>distinguishable parts, solutions including air
  - **<u>heterogeneous</u>** – have visibly distinguishable parts
  - <u>means of physical separation include</u>: filtering, fractional crystallization, distillation, chromatography

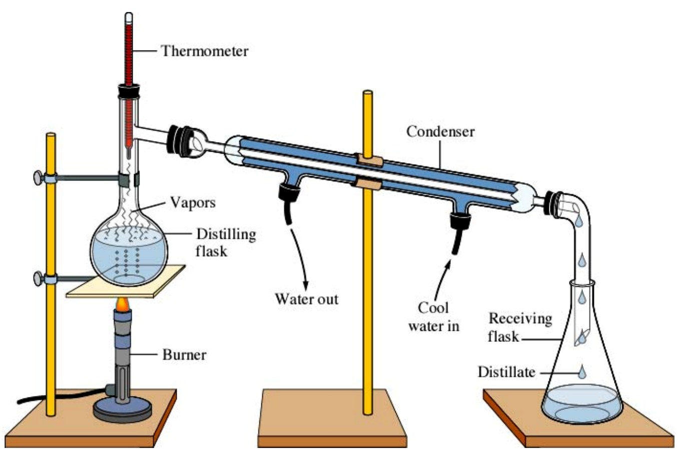
Table 1.5 Densities of Various Common Substances* at 20°C

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

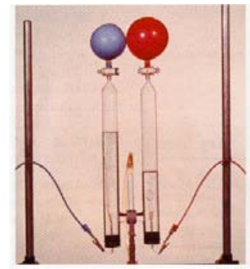
*At 1 atmosphere pressure



Paper chromatograph of ink. (a) A line of the mixture to be separate Is placed at one end of a sheet of porous paper. (b) The paper acts as a wick to draw up the liquid. (c) The component with the weakest attraction for the paper travels faster than those that cling to the paper.



- **<u>Pure substances</u>** – compounds like water, carbon dioxide etc. <u>and</u> elements. Compounds can be separated into elements by **chemical** means
  - electrolysis is a common chemical method for separating compounds into elements.
  - elements can be broken down into atoms
    - which can be broken down into
      - nuclei and electrons
      - p^+, n^0 and e^-
      - quarks



Electrolysis is an example of a chemical change. In this apparatus, water is decomposed to hydrogen gas (filling the red balloon) and Oxygen gas (filling the blue balloon).